

11.5: Concentrations as Conversion Factors

Learning Objective

- Apply concentration units as conversion factors.

Concentration can be a conversion factor between the amount of solute and the amount of solution or solvent (depending on the definition of the concentration unit). As such, concentrations can be useful in a variety of stoichiometry problems. In many cases, it is best to use the original definition of the concentration unit; it is that definition that provides the conversion factor.

A simple example of using a concentration unit as a conversion factor is one in which we use the definition of the concentration unit and rearrange; we can do the calculation again as a unit conversion, rather than as a definition. For example, suppose we ask how many moles of solute are present in 0.108 L of a 0.887 M NaCl solution. Because 0.887 M means 0.887 mol/L, we can use this second expression for the concentration as a conversion factor:

$$0.108 \cancel{L NaCl} \times \frac{0.887 \cancel{mol NaCl}}{1 \cancel{L NaCl}} = 0.0958 \text{ mol NaCl}$$

(There is an understood 1 in the denominator of the conversion factor.) If we used the definition approach, we get the same answer, but now we are using conversion factor skills. Like any other conversion factor that relates two different types of units, the reciprocal of the concentration can be also used as a conversion factor.

Example 11.5.1

Using concentration as a conversion factor, how many liters of 2.35 M CuSO₄ are needed to obtain 4.88 mol of CuSO₄?

Solution

This is a one-step conversion, but the concentration must be written as the reciprocal for the units to work out:

$$4.88 \cancel{\text{mol CuSO}_4} \times \frac{1 \cancel{\text{L}}}{2.35 \cancel{\text{mol}}} = 2.08 \text{ L of solution}$$

Exercise 11.5.1

Using concentration as a conversion factor, how many liters of 0.0444 M CH₂O are needed to obtain 0.0773 mol of CH₂O?

Answer:

1.74 L

Of course, once quantities in moles are available, another conversion can give the mass of the substance, using molar mass as a conversion factor.

Example 11.5.2

What mass of solute is present in 0.765 L of 1.93 M NaOH?

Solution

This is a two-step conversion, first using concentration as a conversion factor to determine the number of moles and then the molar mass of NaOH (40.0 g/mol) to convert to mass:

$$0.765 \cancel{\text{L}} \times \frac{1.93 \cancel{\text{mol NaOH}}}{1 \cancel{\text{L solution}}} \times \frac{40.0 \cancel{\text{g NaOH}}}{1 \cancel{\text{mol NaOH}}} = 59.1 \text{ g NaOH}$$

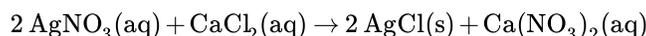
? Exercise 11.5.2

What mass of solute is present in 1.08 L of 0.0578 M H₂SO₄?

Answer

6.12 g

More complex stoichiometry problems using balanced chemical reactions can also use concentrations as conversion factors. For example, suppose the following equation represents a chemical reaction:



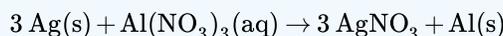
If we wanted to know what volume of 0.555 M CaCl₂ would react with 1.25 mol of AgNO₃, we first use the balanced chemical equation to determine the number of moles of CaCl₂ that would react and then use concentration to convert to liters of solution:

$$1.25 \text{ mol } \cancel{\text{AgNO}_3} \times \frac{1 \text{ mol } \cancel{\text{CaCl}_2}}{2 \text{ mol } \cancel{\text{AgNO}_3}} \times \frac{1 \text{ L solution}}{0.555 \text{ mol } \cancel{\text{CaCl}_2}} = 1.13 \text{ L CaCl}_2$$

This can be extended by starting with the mass of one reactant, instead of moles of a reactant.

✓ Example 11.5.3

What volume of 0.0995 M Al(NO₃)₃ will react with 3.66 g of Ag according to the following chemical equation?



Solution

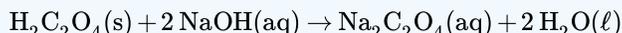
Here, we first must convert the mass of Ag to moles before using the balanced chemical equation and then the definition of molarity as a conversion factor:

$$3.66 \text{ g } \cancel{\text{Ag}} \times \frac{1 \text{ mol } \cancel{\text{Ag}}}{107.97 \text{ g } \cancel{\text{Ag}}} \times \frac{1 \text{ mol } \cancel{\text{Al}(\text{NO}_3)_3}}{3 \text{ mol } \cancel{\text{Ag}}} \times \frac{1 \text{ L solution}}{0.0995 \text{ mol } \cancel{\text{Al}(\text{NO}_3)_3}} = 0.114 \text{ L}$$

The strikeouts show how the units cancel.

? Exercise 11.5.3

What volume of 0.512 M NaOH will react with 17.9 g of H₂C₂O₄(s) according to the following chemical equation?



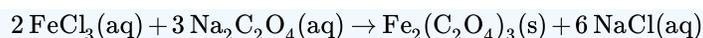
Answer:

0.777 L

We can extend our skills even further by recognizing that we can relate quantities of one solution to quantities of another solution. Knowing the volume and concentration of a solution containing one reactant, we can determine how much of another solution of another reactant will be needed using the balanced chemical equation.

✓ Example 11.5.4

A student takes a precisely measured sample, called an *aliquot*, of 10.00 mL of a solution of FeCl₃. The student carefully adds 0.1074 M Na₂C₂O₄ until all the Fe³⁺(aq) has precipitated as Fe₂(C₂O₄)₃(s). Using a precisely measured tube called a burette, the student finds that 9.04 mL of the Na₂C₂O₄ solution was added to completely precipitate the Fe³⁺(aq). What was the concentration of the FeCl₃ in the original solution? (A precisely measured experiment like this, which is meant to determine the amount of a substance in a sample, is called a *titration*.) The balanced chemical equation is as follows:



Solution

First we need to determine the number of moles of $\text{Na}_2\text{C}_2\text{O}_4$ that reacted. We will convert the volume to liters and then use the concentration of the solution as a conversion factor:

$$9.04 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{0.1074 \text{ mol Na}_2\text{C}_2\text{O}_4}{\text{L}} = 0.000971 \text{ mol Na}_2\text{C}_2\text{O}_4$$

Now we will use the balanced chemical equation to determine the number of moles of $\text{Fe}^{3+}(\text{aq})$ that were present in the initial aliquot:

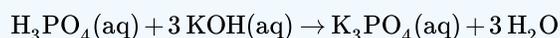
$$0.000971 \text{ mol Na}_2\text{C}_2\text{O}_4 \times \frac{2 \text{ mol FeCl}_3}{3 \text{ mol Na}_2\text{C}_2\text{O}_4} = 0.000647 \text{ mol FeCl}_3$$

Then we determine the concentration of FeCl_3 in the original solution. Converting 10.00 mL into liters (0.01000 L), we use the definition of molarity directly:

$$M = \frac{\text{mol}}{\text{L}} = \frac{0.000647 \text{ mol FeCl}_3}{0.01000 \text{ L}} = 0.0647 \text{ M FeCl}_3$$

? Exercise 11.5.4

A student titrates 25.00 mL of H_3PO_4 with 0.0987 M KOH . She uses 54.06 mL to complete the chemical reaction. What is the concentration of H_3PO_4 ?



Answer:

0.0711 M

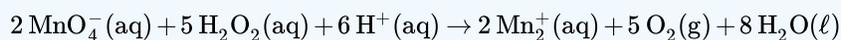


Figure 11.5.1 Titration © Thinkstock. When a student performs a titration, a measured amount of one solution is added to another reactant.

We have used molarity exclusively as the concentration of interest, but that will not always be the case. The next example shows a different concentration unit being used.

✓ Example 11.5.5

H_2O_2 is used to determine the amount of Mn according to this balanced chemical equation:



What mass of 3.00% m/m H_2O_2 solution is needed to react with 0.355 mol of $\text{MnO}_4^- (\text{aq})$?

Solution

Because we are given an initial amount in moles, all we need to do is use the balanced chemical equation to determine the number of moles of H_2O_2 and then convert to find the mass of H_2O_2 . Knowing that the H_2O_2 solution is 3.00% by mass, we can determine the mass of solution needed:

$$0.355 \text{ mol MnO}_4^- \times \frac{5 \text{ mol H}_2\text{O}_2}{2 \text{ mol MnO}_4^-} \times \frac{34.02 \text{ g H}_2\text{O}_2}{\text{mol H}_2\text{O}_2} \times \frac{100 \text{ g solution}}{3 \text{ g H}_2\text{O}_2} = 1006 \text{ g sol}$$

The first conversion factor comes from the balanced chemical equation, the second conversion factor is the molar mass of H_2O_2 , and the third conversion factor comes from the definition of percentage concentration by mass.

? Exercise 11.5.5

Use the balanced chemical reaction for MnO_4^- and H_2O_2 to determine what mass of O_2 is produced if 258 g of 3.00% m/m H_2O_2 is reacted with MnO_4^- .

Answer

7.28 g

Summary

Know how to apply concentration units as conversion factors.

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